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When balancing redox reactions, the overall electronic charge must be balanced in addition to the usual molar ratios of the component reactants and products. This example problem illustrates how to use the half-reaction method to balance a redox reaction in a solution. Balance the following redox reaction in an acidic solution: Cu(s) + HNO3(aq) Cu2+(aq) + NO(g) Step 1: Identify what is being oxidized and what is being reduced. To identify which atoms are being reduced or oxidized, assign oxidation states to each atom of the reaction. For review: Cu(s): Cu = 0HNO3: H = +1, N = +5, O = -6Cu2+: Cu = +2NO(g): N = +2, O = -2 Cu went from oxidation state 0 to +2, losing two electrons. Copper is oxidized by this reaction.N went from oxidation state +5 to +2, gaining three electrons. Nitrogen is reduced by this reaction. Step 2: Break the reaction into two half-reactions: oxidation and reduction. Oxidation: Cu Cu2+ Reduction: HNO3 NO Step 3: Balance each half-reaction by both stoichiometry and electronic charge. This is accomplished by adding substances to the reaction. The only rule is that the only substances you can add must already be in the solution. These include water (H2O), H+ ions (in acidic solutions), OH- ions (in basic solutions) and electrons. Start with the oxidation half-reaction: The half-reaction is already balanced atomically. To balance electronically, two electrons must be added to the product side. Cu Cu2+ + 2 e- Now, balance the reduction reaction. This reaction requires more work. The first step is to balance all atoms except oxygen and hydrogen. HNO3 NO There is only one nitrogen atom on both sides, so nitrogen is already balanced. The second step is to balance the oxygen atoms. This is done by adding water to the side that needs more oxygen. In this case, the reactant side has three oxygens and the product side has only one oxygen. Add two water molecules to the product side. HNO3 NO + 2 H2O The third step is to balance the hydrogen atoms. This is accomplished by adding H+ ions to the side that needs more hydrogen. The reactant side has one hydrogen atom while the product side has four. Add 3 H+ ions to the reactant side. HNO3 + 3 H+ NO + 2 H2O The equation is balanced atomically, but not electrically. The final step is to balance the charge by adding electrons to the more positive side of the reaction. One the reactant side, the overall charge is +3, while the product side is neutral. To counteract the +3 charge, add three electrons to the reactant side. HNO3 + 3 H+ + 3 e- NO + 2 H2O Now the reduction half-equation is balanced. Step 4: Equalize the electron transfer. In redox reactions, the number of electrons gained must equal the number of electrons lost. To accomplish this, each reaction is multiplied by whole numbers to contain the same number of electrons. The oxidation half-reaction has two electrons while the reduction half-reaction has three electrons. The lowest common denominator between them is six electrons. Multiply the oxidation half-reaction by 3 and the reduction half-reaction by 2. 3 Cu 3 Cu2+ + 6 e-2 HNO3 + 6 H+ + 6 e- 2 NO + 4 H2O Step 5: Recombine the half-reactions. This is accomplished by adding the two reactions together. Once they are added, cancel out anything that appears on both sides of the reaction. 3 Cu 3 Cu2+ + 6 e-+ 2 HNO3 + 6 H+ + 6 e- 2 NO + 4 H2O 3 Cu + 2 HNO3 + 6H+ + 6 e- 3 Cu2+ + 2 NO + 4 H2O + 6 e- Both sides have six electrons that can be canceled. 3 Cu + 2 HNO3 + 6 H+ 3 Cu2+ + 2 NO + 4 H2O The complete redox reaction is now balanced. 3 Cu + 2 HNO3 + 6 H+ 3 Cu2+ + 2 NO + 4 H2O To summarize: Identify the oxidation and reduction components of the reaction.Separate the reaction into the oxidation half-reaction and reduction half-reaction.Balance each half-reaction both atomically and electronically.Equalize the electron transfer between oxidation and reduction half-equations.Recombine the half-reactions to form the complete redox reaction. Do you have a redox equation you don't know how to balance? Besides simply balancing the equation in question, these programs will also give you a detailed overview of the entire balancing process with your chosen method. by the ion-electron method In the ion-electron method (also called the half-reaction method), the redox equation is separated into two half-equations - one for oxidation and one for reduction. Each of these half-reactions is balanced separately and then combined to give the balanced redox equation. Enter an equation of a chemical reaction and click 'Submit' (for example: mn2++bio3-+h+=mno4-+bi3+).Rules for typing equations Spaces are irrelevant, for example Cu SO 4 is equal CuSO4 All types of parentheses are correct, for example K3[Fe(CN)6] To enter charge species, just type them as they are, for example Hg2+, Hg22+, or Hg2^2+ To enter the equation sign, you can use either "=" or "-->" or "" symbols. The equation can be written in lowercase letters. If the elements in a chemical formula are properly capitalized, the smart case converter leaves them as you have typed. Why must chemical equations be balanced? A balanced chemical equation accurately describes the quantities of reactants and products in chemical reactions. The Law of Conservation of Mass states that mass is neither created nor destroyed in an ordinary chemical reaction. This means that a chemical equation must have the same number of atoms of each element on both side of the equation. Also the sum of the charges on one side of the equation must be equal to the sum of the charges on the other side. When these two conditions are met, the equation is said to be balanced.Guidelines for balancing redox equations Step 1. Write down the unbalanced equation Step 2. Separate the redox reaction into half-reactions a) Assign oxidation numbers for each atom b) Identify and write out all redox couples in reaction c) Combine these redox couples into two half-reactions Step 3. Balance the atoms in each half reaction a) Balance all other atoms except H and O b) Balance the oxygen atoms with H2O c) Balance the hydrogen atoms with H+ d) In a basic medium, add one OH- to each side for every H+ Step 4. Balance the charge with e- Step 5: Make electron gain equivalent to electron loss in the half-reactions Step 6: Add the half-reactions together Step 7: Simplify the equation Finally, check that the elements and charges are balanced Sample redox equations Ionic vs molecular equation formWhen an equation is written in the molecular form the program will have issues balancing atoms in the half-reactions (Step 3.). This is avoided by writing the equation in the ionic form. Different solutions Chemical reaction with oxidation state changesFor other uses, see Redox (disambiguation).Sodium "gives" one outer electron to fluorine, bonding them to form sodium fluoride. The sodium atom is oxidized, and fluorine is reduced.When a few drops of glycerol (mild reducing agent) are added to powdered potassium permanganate (strong oxidizing agent), a violent redox reaction accompanied by self-ignition starts.Example of a reductionoxidation reaction between sodium and chlorine, with the OIL RIG mnemonic[1]Redox (/rdks/ RED-oks, /ridks/ REE-doks, reductionoxidation[2] or oxidationreduction[3]:150) is a type of chemical reaction in which the oxidation states of the reactants change.[4] Oxidation is the loss of electrons or an increase in the oxidation state, while reduction is the gain of electrons or a decrease in the oxidation state. The oxidation and reduction processes occur simultaneously in the chemical reaction.There are two classes of redox reactions:Electron-transfer Only one (usually) electron flows from the atom, ion, or molecule being oxidized to the atom, ion, or molecule that is reduced. This type of redox reaction is often discussed in terms of redox couples and electrode potentials.Atom transfer An atom transfers from one substrate to another. For example, in the rusting of iron, the oxidation state of iron atoms increases as the iron converts to an oxide, and simultaneously, the oxidation state of oxygen decreases as it accepts electrons released by the iron. Although oxidation reactions are commonly associated with forming oxides, other chemical species can serve the same function.[5] In hydrogenation, bonds like C=C are reduced by transfer of hydrogen atoms."Redox" is a portmanteau of the words "REDuction" and "OXidation." The term "redox" was first used in 1928.[6]Oxidation is a process in which a substance loses electrons. Reduction is a process in which a substance gains electrons.The processes of oxidation and reduction occur simultaneously and cannot occur independently.[5] In redox processes, the reductant transfers electrons to the oxidant. Thus, in the reaction, the reductant or reducing agent loses electrons and is oxidized, and the oxidant or oxidizing agent gains electrons and is reduced. The pair of an oxidizing and reducing agent that is involved in a particular reaction is called a redox pair. A redox couple is a reducing species and its corresponding oxidizing form,[7] e.g., Fe2+/ Fe3+.The oxidation alone and the reduction alone are each called a half-reaction because two half-reactions always occur together to form a whole reaction.[5]In electrochemical reactions the oxidation and reduction processes do occur simultaneously but are separated in space.Main article: Oxidizing agentOxidation originally implied a reaction with oxygen to form an oxide. Later, the term was expanded to encompass substances that accomplished chemical reactions similar to those of oxygen. Ultimately, the meaning was generalized to include all processes involving the loss of electrons or the increase in the oxidation state of a chemical species.[8]:A49 Substances that have the ability to oxidize other substances (cause them to lose electrons) are said to be oxidative or oxidizing, and are known as oxidizing agents, oxidants, or oxidizers. The oxidant removes electrons from another substance, and is thus itself reduced.[8]:A50 Because it "accepts" electrons, the oxidizing agent is also called an electron acceptor. Oxidants are usually chemical substances with elements in high oxidation states[3]:159 (e.g., N2O4, MnO4, CrO3, Cr2O27, OsO4), or else highly electronegative elements (e.g. O2, F2, Cl2, Br2, I2) that can gain extra electrons by oxidizing another substance.[3]:909Oxidizers are oxidants, but the term is mainly reserved for sources of oxygen, particularly in the context of explosions. Nitric acid is a strong oxidizer.[9]The international pictogram for oxidizing chemicalsMain article: Reducing agentSubstances that have the ability to reduce other substances (cause them to gain electrons) are said to be reductive or reducing and are known as reducing agents, reductants, or reducers. The reductant transfers electrons to another substance and is thus itself oxidized.[3]:159 Because it donates electrons, the reducing agent is also called an electron donor. Electron donors can also form charge transfer complexes with electron acceptors. The word reduction originally referred to the loss in weight upon heating a metallic ore such as a metal oxide to extract the metal. In other words, ore was "reduced" to metal.[10] Antoine Lavoisier demonstrated that this loss of weight was due to the loss of oxygen as a gas. Later, scientists realized that the metal atom gains electrons in this process. The meaning of reduction then became generalized to include all processes involving a gain of electrons.[10] Reducing equivalent refers to chemical species which transfer the equivalent of one electron in redox reactions. The term is common in biochemistry.[11] A reducing equivalent can be an electron or a hydrogen atom as a hydride ion.[12]Reductants in chemistry are very diverse. Electropositive elemental metals, such as lithium, sodium, magnesium, iron, zinc, and aluminium, are good reducing agents. These metals donate electrons relatively readily.[13]Hydride transfer reagents, such as NaBH4 and LiAlH4, reduce by atom transfer: they transfer the equivalent of hydride or H. These reagents are widely used in the reduction of carbonyl compounds to alcohols.[14][15] A related method of reduction involves the use of hydrogen gas (H2) as sources of H atoms.[3]:288This section needs expansion. You can help by adding to it. (April 2023)Redox reactions can occur slowly, as in the formation of rust, or rapidly, as in the case of burning fuel. Electron transfer reactions are generally fast, occurring within the time of mixing.[16]The mechanisms of atom-transfer reactions are highly variable because many kinds of atoms can be transferred. Such reactions can also be quite complex, involving many steps. The mechanisms of electron-transfer reactions occur by two distinct pathways, inner sphere electron transfer[17] and outer sphere electron transfer.[18]Analysis of bond energies and ionization energies in water allows calculation of the thermodynamic aspects of redox reactions.[19]Each half-reaction has a standard electrode potential (Eocell), which is equal to the potential difference or voltage at equilibrium under standard conditions of an electrochemical cell in which the cathode reaction is the half-reaction considered, and the anode is a standard hydrogen electrode where hydrogen is oxidized:[20]12H2 H+ + eThe electrode potential of each half-reaction is also known as its reduction potential (Eored), or potential when the half-reaction takes place at a cathode. The reduction potential is a measure of the tendency of the oxidizing agent to be reduced. Its value is zero for H+ + e 12H2 by definition, positive for oxidizing agents stronger than H+ (e.g., +2.866 V for F2) and negative for oxidizing agents that are weaker than H+ (e.g., 0.763V for Zn2+).[8]:873For a redox reaction that takes place in a cell, the potential difference is:Eocell = Eocathode EoanodeHowever, the potential of the reaction at the anode is sometimes expressed as an oxidation potential:Eoox = EoredThe oxidation potential is a measure of the tendency of the reducing agent to be oxidized but does not represent the physical potential at an electrode. With this notation, the cell voltage equation is written with a plus signEocell = Ered(cathode) + Eoox(anode)This section does not cite any sources. Please help improve this section by adding citations to reliable sources. Unsourced material may be challenged and removed. (December 2023) (Learn how and when to remove this message)Illustration of a redox reactionIn the reaction between hydrogen and fluorine, hydrogen is being oxidized and fluorine is being reduced:H2 + F2 2 HFThis spontaneous reaction releases a large amount of energy (542 kJ per 2 g of hydrogen) because two H-F bonds are much stronger than one H-H bond and one F-F bond. This reaction can be analyzed as two half-reactions. The oxidation reaction converts hydrogen to protons:H2 2 H+ + 2 eThe reduction reaction converts fluorine to the fluoride anion:F2 + 2 e 2 FThe half-reactions are combined so that the electrons cancel:H22 H+ + 2 eF2 + 2 e2 FH2 + F22 H+ + 2 FThe protons and fluoride combine to form hydrogen fluoride in a non-redox reaction:2 H+ + 2 F 2 HFThe overall reaction is:H2 + F2 2 HFA redox reaction is the force behind an electrochemical cell like the Galvanic cell pictured. The battery is made out of a zinc electrode in a ZnSO4 solution connected with a wire and a porous disk to a copper electrode in a CuSO4 solution.In this type of reaction, a metal atom in a compound or solution is replaced by an atom of another metal. For example, copper is deposited when zinc metal is placed in a copper(II) sulfate solution:Zn (s) + CuSO4 (aq) ZnSO4 (aq) + Cu (s)In the above reaction, zinc metal displaces the copper(II) ion from the copper sulfate solution, thus liberating free copper metal. The reaction is spontaneous and releases 213 kJ per 65 g of zinc.The ionic equation for this reaction is:Zn + Cu2+ Zn2+ + CuAs two half-reactions, it is seen that the zinc is oxidized:Zn Zn2+ + 2 eAnd the copper is reduced:Cu2+ + 2 e CuThe reduction of nitrate to nitrogen in the presence of an acid (denitrification):2 NO3 + 10 e + 12 H+ N2 + 6 H2OThe combustion of hydrocarbons, such as in an internal combustion engine, produces water, carbon dioxide, some partially oxidized forms such as carbon monoxide, and heat energy. Complete oxidation of materials containing carbon produces carbon dioxide.The stepwise oxidation of a hydrocarbon by oxygen, in organic chemistry, produces water and, successively: an alcohol, an aldehyde or a ketone, a carboxylic acid, and then a peroxide.Oxides, such as iron(III) oxide or rust, which consists of hydrated iron(III) oxides Fe2O3nH2O and iron(III) oxide-hydroxide (FeO(OH), Fe(OH)3), form when oxygen combines with other elements.Iron rusting in pyrite cubesThe term corrosion refers to the electrochemical oxidation of metals in reaction with an oxidant such as oxygen. Rusting, the formation of iron oxides, is a well-known example of electrochemical corrosion: it forms as a result of the oxidation of iron metal. Common rust often refers to iron(III) oxide, formed in the following chemical reaction:4 Fe + 3 O2 2 Fe2O3The oxidation of iron(II) to iron(III) by hydrogen peroxide in the presence of an acid:Fe2+ Fe3+ + eH2O2 + 2 e 2 OHHere the overall equation involves adding the reduction equation to twice the oxidation equation, so that the electrons cancel:2 Fe2+ + H2O2 + 2 H+ 2 Fe3+ + 2 H2OA disproportionation reaction is one in which a single substance is both oxidized and reduced. For example, thiosulfate ion with sulfur in oxidation state +2 can react in the presence of acid to form elemental sulfur (oxidation state 0) and sulfur dioxide (oxidation state +4).S2O23 + 2 H+ S + SO2 + H2OThus one sulfur atom is reduced from +2 to 0, while the other is oxidized from +2 to +4.[8]:176Cathodic protection is a technique used to control the corrosion of a metal surface by making it the cathode of an electrochemical cell. A simple method of protection connects protected metal to a more easily corroded "sacrificial anode" to act as the anode. The sacrificial metal, instead of the protected metal, then corrodes. Oxidation is used in a wide variety of industries, such as in the production of cleaning products and oxidizing ammonia to produce nitric acid.[citation needed]Redox reactions are the foundation of electrochemical cells, which can generate electrical energy or support electrosynthesis. Metal ores often contain metals in oxidized states, such as oxides or sulfides, from which the pure metals are extracted by smelting at high temperatures in the presence of a reducing agent. The process of electroplating uses redox reactions to coat objects with a thin layer of a material, as in chrome-plated automotive parts, silver plating cutlery, galvanization and gold-plated jewelry.[citation needed]This section does not cite any sources. Please help improve this section by adding citations to reliable sources. Unsourced material may be challenged and removed. (December 2023) (Learn how and when to remove this message)Top: ascorbic acid (reduced form of Vitamin C)Bottom: dehydroascorbic acid (oxidized form of Vitamin C)Enzymatic browning is an example of a redox reaction that takes place in most fruits and vegetables.Many essential biological processes involve redox reactions. Before some of these processes can begin, iron must be assimilated from the environment.[21]Cellular respiration, for instance, is the oxidation of glucose (C6H12O6) to CO2 and the reduction of oxygen to water. The summary equation for cellular respiration is:C6H12O6 + 6 O2 6 CO2 + 6 H2O + EnergyThe process of cellular respiration also depends heavily on the reduction of NAD+ to NADH and the reverse reaction (the oxidation of NADH to NAD+). Photosynthesis and cellular respiration are complementary, but photosynthesis is not the reverse of the redox reaction in cellular respiration:6 CO2 + 6 H2O + light energy C6H12O6 + 6 O2Biological energy is frequently stored and released using redox reactions. Photosynthesis involves the reduction of carbon dioxide into sugars and the oxidation of water into molecular oxygen. The reverse reaction, respiration, oxidizes sugars to produce carbon dioxide and water. As intermediate steps, the reduced carbon compounds are used to reduce nicotinamide adenine dinucleotide (NAD+) to NADH, which then contributes to the creation of a proton gradient, which drives the synthesis of adenosine triphosphate (ATP) and is maintained by the reduction of oxygen. In animal cells, mitochondria perform similar functions.See also: Membrane potentialThe term redox state is often used to describe the balance of GSH/GSSG, NAD+/NADH and NADP+/NADPH in a biological system such as a cell or organ. The redox state is reflected in the balance of several sets of metabolites (e.g., lactate and pyruvate, beta-hydroxybutyrate and acetoacetate), whose interconversion is dependent on these ratios. Redox mechanisms also control some cellular processes. Redox proteins and their genes must be co-located for redox regulation according to the CoRR hypothesis for the function of DNA in mitochondria and chloroplasts.Wide varieties of aromatic compounds are enzymatically reduced to form free radicals that contain one more electron than their parent compounds. In general, the electron donor is any of a wide variety of flavoenzymes and their coenzymes. Once formed, these anion free radicals reduce molecular oxygen to superoxide and regenerate the unchanged parent compound. The net reaction is the oxidation of the flavoenzyme's coenzymes and the reduction of molecular oxygen to form superoxide. This catalytic behavior has been described as a futile cycle or redox cycling.Blast furnaces of Tinec Iron and Steel Works, Czech RepublicMinerals are generally oxidized derivatives of metals. Iron is mined as ores such as magnetite (Fe3O4) and hematite (Fe2O3). Titanium is mined as its dioxide, usually in the form of rutile (TiO2). These oxides must be reduced to obtain the corresponding metals, often achieved by heating these oxides with carbon or carbon monoxide as reducing agents. Blast furnaces are the reactors where iron oxides and coke (a form of carbon) are combined to produce molten iron. The main chemical reaction producing the molten iron is:[22]Fe2O3 + 3 CO 2 Fe + 3 CO2Electron transfer reactions are central to myriad processes and properties in soils, and redox potential, quantified as Eh (platinum electrode potential (voltage) relative to the standard hydrogen electrode) or pe (analogous to pH as log electron activity), is a master variable, along with pH, that controls and is governed by chemical reactions and biological processes. Early theoretical research with applications to flooded soils and paddy rice production was seminal for subsequent work on thermodynamic aspects of redox and plant root growth in soils.[23] Later work built on this foundation, and expanded it for understanding redox reactions related to heavy metal oxidation state changes, pedogenesis and morphology, organic compound degradation and formation, free radical chemistry, wetland delineation, soil remediation, and various methodological approaches for characterizing the redox status of soils.[24][25]Main article: List of chemistry mnemonicsThe key terms involved in redox can be confusing.[26][27] For example, a reagent that is oxidized loses electrons; however, that reagent is referred to as the reducing agent. Likewise, a reagent that is reduced gains electrons and is referred to as the oxidizing agent.[28] These mnemonics are commonly used by students to help memorise the terminology:[29]"OIL RIG" oxidation is loss of electrons, reduction is gain of electrons[26][27][28][29]"LEO the lion says GER [grr]" loss of electrons is oxidation, gain of electrons is reduction[26][27][28][29]"LEORA says GEROA" the loss of electrons is called oxidation (reducing agent); the gain of electrons is called reduction (oxidizing agent).[28]"RED CAT" and "AN OX", or "AnOx RedCat" ("an ox-red cat") reduction occurs at the cathode and the anode is for oxidation"RED CAT gains what AN OX loses" reduction at the cathode gains (electrons) what anode oxidation loses (electrons)"PANIC" Positive Anode and Negative is Cathode. This applies to electrolytic cells which release stored electricity, and can be recharged with electricity. PANIC does not apply to cells that can be recharged with redox materials. These galvanic or voltaic cells, such as fuel cells, produce electricity from internal redox reactions. Here, the positive electrode is the cathode and the negative is the anode.Anaerobic respirationBessemer processBioremediationCalvin cycleChemical equationChemical looping combustionCitric acid cycleElectrochemical seriesElectrochemistryElectrolysisElectron equivalentElectron transport chainElectrosynthesisGalvanic cellHydrogenationMembrane potentialMicrobial fuel cellMurburn conceptNucleophilic abstractionOrganic redox reactionOxidative addition and reductive eliminationOxidative phosphorylationPartial oxidationPro-oxidantRedox gradientRedox potentialRedox therapyReducing agentReducing atmosphereReduction potentialThermic reactionTransmetalationSulfur cycle^ "Metals". Bitesize. BBC. Archived from the original on November 3, 2022.^ "redox definition of redox in English | Oxford Dictionaries". Oxford Dictionaries | English. Archived from the original on October 1, 2017. 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ISBN978-0-8412-2652-4.Wikiquote has quotations related to Redox.Wikimedia Commons has media related to Redox reactions.Retrieved from " Share copy and redistribute the material in any medium or format for any purpose, even commercially. Adapt remix, transform, and build upon the material for any purpose, even commercially. The licensor cannot revoke these freedoms as long as you follow the license terms. Attribution You must give appropriate credit , provide a link to the license, and indicate if changes were made . You may do so in any reasonable manner, but not in any way that suggests the licensor endorses you or your use. ShareAlike If you remix, transform, or build upon the material, you must distribute your contributions under the same license as the original. No additional restrictions You may not apply legal terms or technological measures that legally restrict others from doing anything the license permits. You do not have to comply with the license for elements of the material in the public domain or where your use is permitted by an applicable exception or limitation . No warranties are given. The license may not give you all of the permissions necessary for your intended use. For example, other rights such as publicity, privacy, or moral rights may limit how you use the material. Do you have a redox equation you don't know how to balance? Besides simply balancing the equation in question, these programs will also give you a detailed overview of the entire balancing process with your chosen method. by the ion-electron method In the ion-electron method (also called the half-reaction method), the redox equation is separated into two half-equations - one for oxidation and one for reduction. Each of these half-reactions is balanced separately and then combined to give the balanced redox equation. Enter an equation of a chemical reaction and click 'Submit' (for example: mn2++bio3-+h+=mno4-+bi3+).Rules for typing equations Spaces are irrelevant, for example Cu SO 4 is equal CuSO4 All types of parentheses are correct, for example K3[Fe(CN)6] To enter charge species, just type them as they are, for example Hg2+, Hg22+, or Hg2^2+ To enter the equation sign, you can use either "=" or "-->" or "" symbols. The equation can be written in lowercase letters. If the elements in a chemical formula are properly capitalized, the smart case converter leaves them as you have typed. Why must chemical equations be balanced? A balanced chemical equation accurately describes the quantities of reactants and products in chemical reactions. The Law of Conservation of Mass states that mass is neither created nor destroyed in an ordinary chemical reaction. This means that a chemical equation must have the same number of atoms of each element on both side of the equation. Also the sum of the charges on one side of the equation must be equal to the sum of the charges on the other side. When these two conditions are met, the equation is said to be balanced.Guidelines for balancing redox equations Step 1. Write down the unbalanced equation Step 2. Separate the redox reaction into half-reactions a) Assign oxidation numbers for each atom b) Identify and write out all redox couples in reaction c) Combine these redox couples into two half-reactions Step 3. Balance the atoms in each half reaction a) Balance all other atoms except H and O b) Balance the oxygen atoms with H2O c) Balance the hydrogen atoms with H+ d) In a basic medium, add one OH- to each side for every H+ Step 4. Balance the charge with e- Step 5: Make electron gain equivalent to electron loss in the half-reactions Step 6: Add the half-reactions together Step 7: Simplify the equation Finally, check that the elements and charges are balanced Sample redox equations Ionic vs molecular equation formWhen an equation is written in the molecular form the program will have issues balancing atoms in the half-reactions (Step 3.). This is avoided by writing the equation in the ionic form. Different solutions Oxidation Number / State Method For Balancing Redox Reactions This method is based on the principle that the number of electrons lost in oxidation must be equal to the number of electrons gained in reduction. The steps to be followed are : Write the equation (if it is not complete, then complete it) representing the chemical changes. By knowing oxidation numbers of elements, identify which atom(s) is(are) undergoing oxidation and reduction. Write down separate equations for oxidation and reduction. Add respective electrons on the right of oxidation reaction and on the left of reduction reaction. Care must be taken to ensure that the net charge on both the sides of the equation is same. Multiply the oxidation and reduction reactions by suitable numbers to make the number of electrons lost in oxidation reactions equal to the number of electrons gained in reduction reactions. Transfer the coefficient of the oxidizing and reducing agents and their products to the main equation. By inspection, arrive at the co-efficients of the species not undergoing oxidation or reduction. Balance the equation : KMnO4+ H2SO4+ FeSO4K2SO4+ MnSO4+ Fe2(SO4)3+ H2O Solution: Mn+7Mn+2; Fe+2 Fe2+3 Mn+7+ 5e- Mn+2(1)2; 2Fe+2Fe2+3+ 2e-(2)5 2Mn+7+ 10e-2Mn+2; 10Fe2+25Fe2+3+ 10e-. Therefore coefficient of KMnO4is 2, that of MnSO4is 2, that of FeSO4is 10 and that of Fe2(SO4) is 5. Therefore 2KMnO4+ 10FeSO42MnSO4+ 5Fe2(SO4)3 Coefficient of K2SO4should be 1. 2KMnO4+ 10 FeSO4K2SO4+ 2MnSO4+ 5Fe2(SO4)3 To balance SO42-so that left hand side must have 8H2SO4, thus H2O also becomes 8H2O. 2KMnO4+ 10 FeSO4+ 8H2SO4K2SO4+ 2Mn(SO4) + 5Fe2(SO4)3+ 8H2O _____ Question 2: Balance the equation : MnO4-+ C2O42 + H+CO2+ Mn+2+ H2O Solution: Mn+7 Mn+2; C2C+4 5e-+ Mn+7Mn+2 (1) 2; C22C+4+ 2e- (2) 5 10e-+ 2Mn+72Mn+2; 5C210C+4+ 10e- Therefore 2MnO4-+ 5C2O4-210CO2-+ 2Mn+2 There must be 8H2O on the right to balance O. Therefore there must be 10H+to balance H. 2MnO4-+ 5C2O4-2+ 16H+10CO2+ 2Mn+2+ 8H2O Half-Reaction or Ion-Electron MethodFor Balancing Redox Reactions This method involves the following steps : Divide the complete equation into two half reactions, one representing oxidation and the other reduction. Balance the atoms in each half reaction separately according to the following steps: First of all balance the atoms other than H and O. In a reaction taking place in acidic or neutral medium, oxygen atoms are balanced by adding molecules of water to the side deficient in oxygen atoms while hydrogen atoms are balanced by adding H+ ions to the other side deficient in hydrogen atoms. On the other hand, in alkaline medium (OH-), every excess of oxygen atom on one side is balanced by adding one H2O to the same side and 2OH- to the other side. In case hydrogen is still unbalanced, then balance by adding one OH-, for every excess of H atom on the same side as the excess and one H2O on the other side. Equalize the charge on both sides by adding a suitable number of electrons to the side deficient in negative charge. Multiply the two half reactions by suitable integers so that the total number of electrons gained in one half reaction is equal to the number of electrons lost in the other half reaction. Watch this Video for more reference Add the two balanced half equations and cancel any term common to both sides. Balance the redox reaction Al + NO3- Al(OH)4-+NH3in alkaline medium. Solution: Dividing the equation in two half reactions NO3-NH3 (Oxidation Half Reaction) Al + 4OH-Al(OH)4-(Reduction Half Reaction) All the atoms other than O & H are balanced (N and Al in this case which are already balanced). NO3-+ H2O NH3+ OH- Al + 4OH-Al(OH)4- Then O and H are balanced using OH- and H2O NO3-+ 6H2O NH3+ 9OH- Al + 4 OH- Al(OH)4- Electrons need to be added in order to balance the charge. 8 e- + NO3-+ 6 H2O NH3+ 9 OH- Al + 4 OH- Al(OH)4-+ 3 e- Loss and gain of electrons need to be made equal: For this, we need to multiply first equation by 3 and second equation by 8 Adding both equations we get the balanced redox equation 3NO3-+ 18H2O+ 8Al + 5 OH- 8Al(OH)4-+3NH3 _____ Question 2: Balance the equation Cr2O72-(g) + SO2(aq) Cr3+ + SO42-(aq) (in acidic medium) Solution: Dividing the equation in two half reactions SO2SO42- (Oxidation half reaction) Cr2O72-Cr3+ (Reduction half reaction) Balance the atoms other than O and H SO2SO42- Cr2O72-2Cr3+ O and H are balanced using H+ and H2O SO2 + 2H2OSO42- +4H+ Cr2O72-+14H+2Cr3+ +7H2O Electrons need to be added to balance the charge. SO2 + 2H2OSO42- +4H++2e-.....(2) Cr2O72-+14H+ +6e-2Cr3+ +7H2O.....(1) Loss and gain of electrons need to be made equal.For this, we need to multiply equation (2) by 3 Adding both equations to get the balanced redox reaction Cr2O72-+2H+ +3SO22Cr3+ +7H2O Question 1: Which of the following equations represents a balanced redox reaction? a. KMnO4+ H2SO4+ FeSO4K2SO4+ MnSO4+ Fe2(SO4)3+ H2Ob. 2KMnO4+ 10 FeSO4+ 8H2SO4K2SO4+ 2Mn(SO4) + 5Fe2(SO4)3+ 8H2Oc. Cr2O72-(g) + SO2(aq) Cr3++ SO42-(aq)d. Al + NO3- Al(OH)4-+NH3 Question 2:Ion-electron method for balancing Redox reactions is also known as a. Oxidation number method b. State method c. Half reaction method d. Half life method Question 3:In a reaction taking place in acidic or neutral medium, oxygen atoms are balanced by adding a.H2O & H+ b. H2O & OH- c. H+ & OH- d. H2O only Question 4:What is the number of electrons being transferredin the redox reaction2KMnO4+ 10 FeSO4+ 8H2SO4K2SO4+ 2Mn(SO4) + 5Fe2(SO4)3+ 8H2O? a. Zero b. One c. Five d. Ten Related Resources To read more,Buy study materials of Redox Reactionscomprising study notes, revision notes, video lectures, previous year solved questions etc. Also browse for more study materials on Chemistry here.Its time to prepare for your A-level chemistry exam. Dont be daunted by the volume of material you need to go through or anything you may not know 100% yet, thats where we at A-level Chemistry come in with loads of revision notes of your coursework no matter if its AQA, OCR or EdExcel. Weve got revision notes that will help you familiarize yourself with the major themes and topics that tend to come up in the exams so youll feel prepared and confident before you enter the